OXIDATION AND REDUCTION (REDOX) REACTIONS

TRADITIONAL DEFINITION

OXIDATION
- Gain of oxygen.
- Loss of hydrogen.

O_2

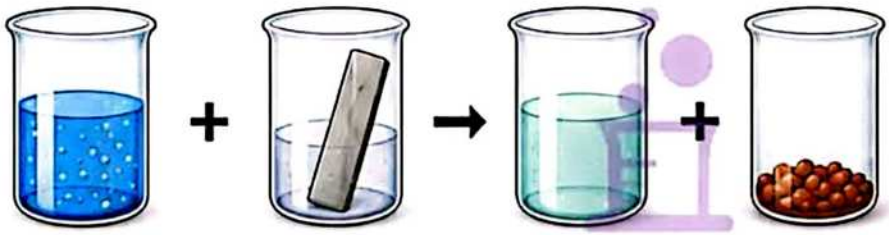
REDUCTION
- Loss of oxygen.
- Gain of hydrogen.

H_2

ELECTRONIC (MODERN) DEFINITION

Example: Zinc Reacting with Copper(II) Sulfate

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$



Copper(II) Sulfate Solution ($CuSO_4$) | Zinc Strip (Zn) | Zinc Sulfate Solution ($ZnSO_4$) | Copper (Cu)

What Happens?
- Zinc loses two electrons → It is **oxidized.**
- Copper(II) ions gain two electrons → They are **reduced.**

HALF-REACTIONS

Oxidation:
$$Zn \rightarrow Zn^{2+} + 2e^-$$

Reduction:
$$Cu^{2+} + 2e^- \rightarrow Cu$$

KEY IDEA
Oxidation and reduction are not limited to oxygen and hydrogen. They also involve the transfer of electrons between substances.

IMPORTANT DEFINITIONS

OXIDATION
Loss of one or more electrons by an atom or ion.

$Zn \xrightarrow{-2e^-} Zn^{2+}$

REDUCTION
Gain of one or more electrons by an atom or ion.

$Cu^{2+} \xrightarrow{+2e^-} Cu$

REDUCING AGENT
A substance that loses electrons and causes another substance to be reduced. → It is **oxidized** during the reaction.

Example: Zinc (Zn)

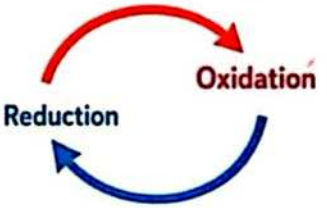
$Zn \rightarrow Zn^{2+} + 2e^-$

OXIDIZING AGENT
A substance that gains electrons and causes another substance to be oxidized. → It is **reduced** during the reaction.

Example: Copper(II) ion (Cu^{2+})

$Cu^{2+} + 2e^- \rightarrow Cu$

IMPORTANT NOTE



- ✓ Oxidation and reduction always occur together in the same reaction.
- ✓ Oxidation cannot occur without reduction.
- ✓ Reduction cannot occur without oxidation.

These paired processes are called Redox Reactions.

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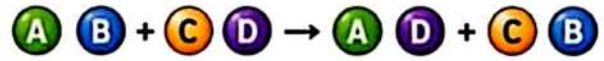
4. DOUBLE DISPLACEMENT REACTIONS

DEFINITION

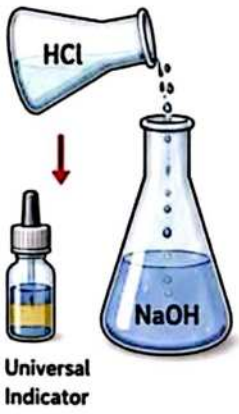
A Double Displacement Reaction occurs when two ionic compounds, usually in aqueous solution, exchange their positive and negative ions, forming two new compounds.

GENERAL FORM



1 ACID-BASE NEUTRALIZATION REACTION

Hydrochloric acid (HCl) reacts with sodium hydroxide (NaOH) to produce salt and water.



PRACTICAL EXPERIMENT

1. Add a few drops of Universal Indicator to a sodium hydroxide (NaOH) solution.
2. Slowly add hydrochloric acid (HCl) while stirring.
3. Continue adding acid until the indicator changes color, indicating the neutralization point.

RESULT

A neutral solution is formed containing:

- Sodium Chloride (NaCl)
- Water (H_2O)

CONCLUSION

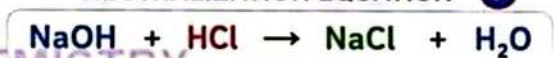
Acids react with alkalis in a neutralization reaction, producing salt and water.

UNIVERSAL INDICATOR NOTES

The Universal Indicator changes color depending on the solution's pH.

INDICATOR COLORS

NEUTRALIZATION EQUATION



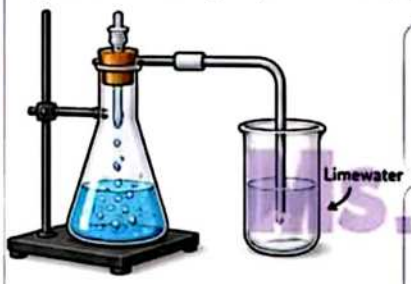
STRONG ACIDS vs WEAK ACIDS

Comparison	Strong Acid	Weak Acid
Ionization in Water	Completely ionizes	Partially ionizes
H^+ Concentration	High	Low
Examples	HCl, HNO_3, H_2SO_4	CH_3COOH, H_2CO_3

2 STRONG ACID + SALT OF A WEAK ACID

Hydrochloric acid reacts with sodium carbonate to produce:

- Sodium Chloride (NaCl)
- Water (H_2O)
- Carbon Dioxide (CO_2)



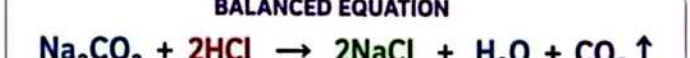
DETECTION OF CO_2

Carbon dioxide turns limewater milky.

PRACTICAL APPLICATION

Use a flask, delivery tube, and limewater to detect carbon dioxide gas.

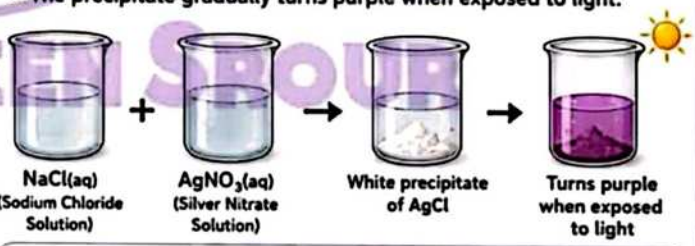
BALANCED EQUATION



3 PRECIPITATION REACTION (BETWEEN TWO SALT SOLUTIONS)

When silver nitrate solution is mixed with sodium chloride solution:

- A white precipitate of silver chloride (AgCl) forms.
- The precipitate gradually turns purple when exposed to light.



CHEMICAL EQUATION

REAL-LIFE APPLICATION: WATER SOFTENING

WHAT IS HARD WATER?	THE PROBLEM	THE SOLUTION	BENEFIT
Hard water contains high concentrations of: • Calcium ions (Ca^{2+}) • Magnesium ions (Mg^{2+})	These ions form solid deposits (scale) inside pipes and heating systems, reducing efficiency and increasing energy consumption.	Add Sodium Carbonate (Na_2CO_3) to precipitate calcium and magnesium ions as insoluble solids.	Removing hardness prevents scale formation, improves the efficiency of appliances and pipes, and extends their service life.

$Ca^{2+} + CO_3^{2-} \rightarrow CaCO_3\downarrow$

$Mg^{2+} + CO_3^{2-} \rightarrow MgCO_3\downarrow$

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METAL DISPLACING HYDROGEN FROM WATER

1 HIGHLY REACTIVE METALS

Metals that are above hydrogen in the reactivity series (such as Na) react with water and displace hydrogen, releasing hydrogen gas (H_2).

Example: Na, Mg > H > Cu ....

2 SODIUM EXPERIMENT

Place a small piece of sodium (about the size of a grain of rice), dried with kerosene, on the surface of water.

3 CHEMICAL EQUATION

$$2Na + 2H_2O \rightarrow 2NaOH + H_2\uparrow + Heat$$

4 HAZARD

This reaction is highly exothermic and may ignite or explode due to the rapid production of hydrogen gas.

5 SAFETY PRECAUTION

Never use water to extinguish a sodium fire.
Use dry sand instead.

6 REACTIVITY COMPARISON

Na

Sodium (Na) reacts immediately with cold water.

Mg

Magnesium (Mg) reacts only with hot water or steam.

Conclusion: Magnesium is less reactive than sodium.

METAL DISPLACING ANOTHER METAL FROM ITS SALT SOLUTION

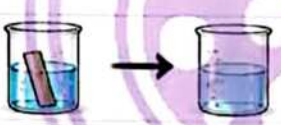
1 PRINCIPLE

A more reactive metal displaces a less reactive metal from its salt solution.

Example: Zn → Cu

2 ZINC AND COPPER EXPERIMENT

Zn



Immerse a zinc strip (Zn) in a copper(II) sulfate solution ($CuSO_4$).

3 OBSERVATION

- The blue color gradually fades.
- A reddish-brown layer of copper is deposited on the zinc strip.

4 CONCLUSION

Zinc is more reactive than copper, so it displaces copper from its salt solution.

$$Zn > Cu$$

5 CHEMICAL EQUATION

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

6 GENERAL RULE

The metal that is higher in the reactivity series displaces the metal below it from its salt solution.

PRACTICAL APPLICATION
Cleaning Tarnished Silver

Aluminum foil is used to clean tarnished silver because aluminum is more reactive than silver.

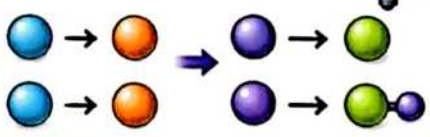
It removes the black silver sulfide (Ag_2S) coating from the silver surface.

HOW THE CLEANING PROCESS WORKS

Place the tarnished silver in hot water with:
- Aluminum foil
- Sodium bicarbonate (baking soda)

Since aluminum is more reactive than silver, it removes sulfur from the silver and restores its shine.

KEY IDEA



Chemical reactions are not random. They follow predictable reaction patterns, such as:
- Combination (Synthesis)
- Decomposition
- Single Displacement
- Double Displacement

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3. SINGLE DISPLACEMENT REACTIONS

DEFINITION

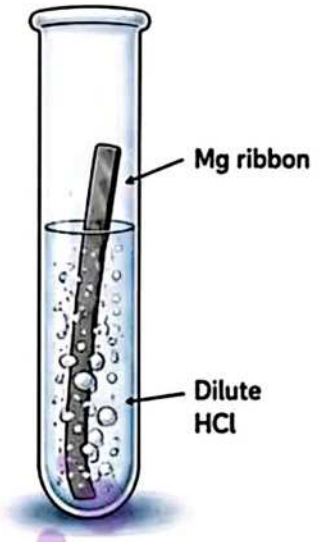
A Single Displacement Reaction is a chemical reaction in which a more reactive element replaces a less reactive element in a compound.

GENERAL EQUATION

$$A + BC \rightarrow AC + B$$

A	BC	AC	B
(More reactive element)	(Compound)	(New compound)	(Displaced element)

EXPERIMENT: MAGNESIUM DISPLACES HYDROGEN FROM HYDROCHLORIC ACID

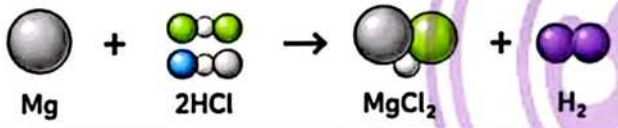


- **MATERIALS:**
  - Magnesium ribbon (Mg)
  - Dilute Hydrochloric Acid (HCl)
  - Test tube
- **PROCEDURE:** Place a magnesium ribbon into a test tube containing dilute hydrochloric acid.
- **OBSERVATION:** Gas bubbles are produced (Hydrogen gas).
- **RESULT:** Hydrogen gas is released.
- **TEST:** Bring a burning splint near the mouth of the test tube. A "pop" sound confirms the presence of hydrogen.

CHEMICAL EQUATION

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \uparrow$$

- Magnesium replaces hydrogen in hydrochloric acid.
- Products:
  - Magnesium Chloride ($MgCl_2$)
  - Hydrogen Gas (H_2)



Mg	2HCl	$MgCl_2$	H_2

IMPORTANT CONCLUSION

- Active metals such as Magnesium (Mg) replace hydrogen from dilute acids.
- Hydrogen gas is released during the reaction.
- The faster the bubbles appear, the more reactive the metal.
- The reaction rate depends on the metal's chemical reactivity.

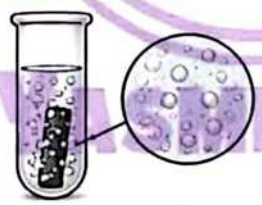
PRACTICAL ACTIVITY (COMPARISON OF THREE METALS)

React the following metals with dilute hydrochloric acid.

RESULT:

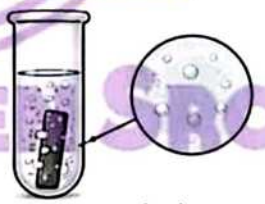
- Zinc (Zn): Reacts rapidly with many bubbles.
- Iron (Fe): Reacts more slowly than zinc.
- Copper (Cu): No reaction.

Zn + HCl



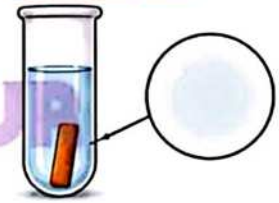
Reacts rapidly with many bubbles.

Fe + HCl



Reacts more slowly than zinc.

Cu + HCl



No reaction.

REACTIVITY SERIES (THEORETICAL RULE)

Metals arranged from most reactive to least reactive

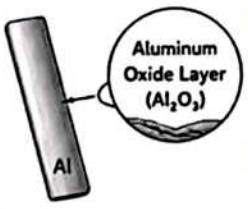
K Na Ca Mg Al Zn Fe H Cu Ag Au

React with dilute acids and release hydrogen gas.

Hydrogen (Reference)

Do not react with dilute acids.

IMPORTANT EXCEPTION: ALUMINUM (Al)



1. The oxide layer protects the metal.
2. The acid gradually dissolves the oxide layer.
3. Once the oxide layer is removed, aluminum reacts rapidly.

Although aluminum is more reactive than zinc, it reacts slowly at first.

FINAL SUMMARY

The chemical reactivity of a metal determines whether it reacts with dilute acids. However, surface factors, such as the presence of an oxide layer, may delay the start of the reaction even for highly reactive metals.

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CHEMICAL REACTIONS

CONCEPT OF A CHEMICAL REACTION

- Breaking chemical bonds in the reactants and forming new bonds in the products.
- Produces new substances with different properties.

EVIDENCE OF A CHEMICAL REACTION

- Color or odor change
- Gas evolution
- Formation of a precipitate
- Release of heat or light

MAIN TYPES OF CHEMICAL REACTIONS

1. Combination (Synthesis) Reactions
2. Decomposition Reactions
3. Single Displacement Reactions
4. Double Displacement Reactions

1 COMBINATION (SYNTHESIS) REACTIONS

GENERAL EQUATION

$$A + B \rightarrow AB$$

EXAMPLE: BURNING MAGNESIUM IN OXYGEN

$$2Mg + O_2 \rightarrow 2MgO$$

EXPERIMENT: HEATING MAGNESIUM

- **MATERIALS:**
  Bunsen burner
  Metal tongs
  Magnesium ribbon (Mg)
- **OBSERVATION:**
  Magnesium burns with a bright white flame.
- **CONCLUSION:**
  Magnesium combines with oxygen to form Magnesium Oxide (MgO).

REACTION OF METALS WITH OXYGEN (COMPARISON)

Observation	Equation	Metal	Reaction
Most reactive	$2Mg + O_2 \rightarrow 2MgO$	Magnesium (Mg)	Burns with a bright white flame
Moderately reactive	$3Fe + 2O_2 \rightarrow Fe_3O_4$	Iron (Fe)	Burns producing sparks
Least reactive	$2Cu + O_2 \rightarrow 2CuO$	Copper (Cu)	Does not burn; forms a black coating

IMPORTANT CONCLUSION

Most metals combine with oxygen when heated, but their chemical reactivity differs:

$$Mg > Fe > Cu$$

Magnesium is more reactive than iron, and iron is more reactive than copper.

SAFETY PRECAUTIONS

Magnesium emits an intense bright light and ultraviolet (UV) radiation when burning.
Always wear safety goggles during the experiment.

2 DECOMPOSITION REACTIONS (COMPOUND BREAKDOWN)

EXAMPLE 1

Decomposition of Mercury(II) Oxide

$$2HgO \rightarrow 2Hg + O_2$$

(Heating)

Mercury(II) Oxide (HgO) → Mercury (Hg) + Oxygen (O_2)

EXAMPLE 2

Decomposition of Calcium Carbonate

$$CaCO_3 \rightarrow CaO + CO_2$$

(Heating)

Calcium Carbonate ($CaCO_3$) → Calcium Oxide (CaO) + Carbon Dioxide (CO_2)

SUMMARY

Heating Calcium Carbonate ($CaCO_3$) produces:

- Calcium Oxide (CaO)
- Carbon Dioxide (CO_2)

Equation:

$CaCO_3$ —Heat→ $CaO + CO_2$

$$CaCO_3 \rightarrow CaO + CO_2$$

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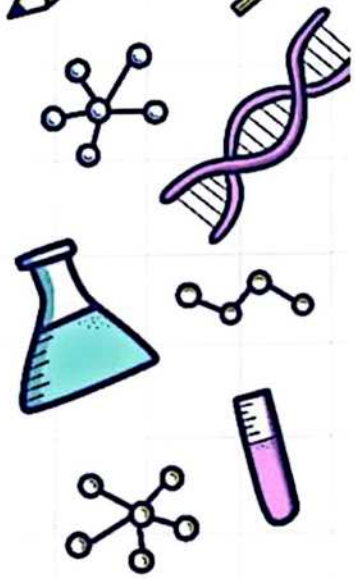

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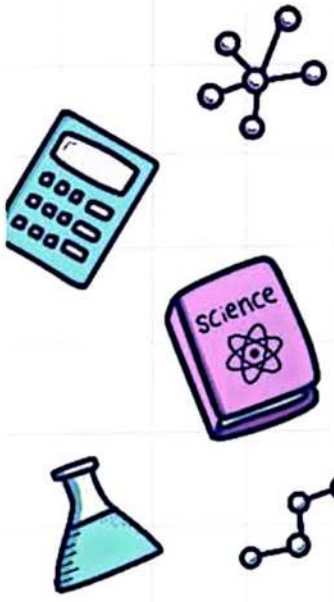

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